CHEMICAL EQUILIBRIUM 14 MAY 2013

Lesson Description

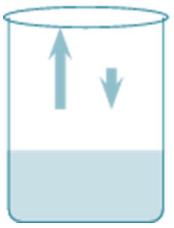
In this lesson, we:

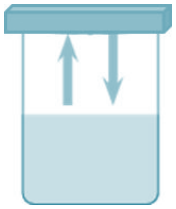
- Define phase equilibrium
- Define and discuss chemical equilibrium
- State Le Chatelier's principle
- Explain changes in chemical equilibrium

Key Concepts

Phase Equilibrium

When water is in an open system, such as an open beaker, water particles are able to escape. Eventually all the water will evaporate. The water particles will not condense as the system is open so no equilibrium can be established.





However if the beaker is sealed off, the water will eventually get to a place when a phase equilibrium will be formed. This means that as quickly as the water evaporates some of the water will condense.

So if a person was watching the beaker, they would see the water level drop for a period of time and then it would appear as if the water has stopped evaporating as the water level will remain constant.

A dynamic phase equilibrium has been reached with the water. This means that the rate of evaporation is equal to the rate of condensation and both processes are happening continuously.

Phase equilibrium can be represented as follows: $Liquid \rightleftharpoons vapour$

Phase equilibrium is an equilibrium that is established with a physical process as phase change is a physical change.

It is possible to establish a chemical equilibrium as well.

Chemical Equilibrium

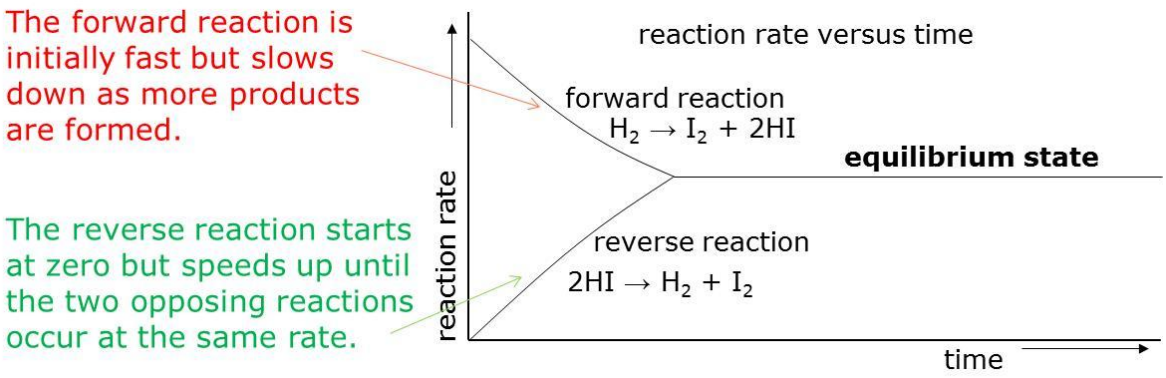
Many chemical reactions are spontaneously reversible if they are in a closed system.

For example: $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$

When the hydrogen and Iodine are first placed into the reaction container, they will quickly form hydrogen iodide. This means that the forward reaction will be fast at the beginning. However, as the hydrogen iodide is formed, so hydrogen iodide molecules will collide and react to form hydrogen and iodine, thus the reverse reaction starts to take place.

Eventually the rate at which hydrogen iodide is being formed will be equal to the rate at which hydrogen and iodine is formed. So the forward rate is equal to the reverse rate.

This can be represented on a rate graph, as shown below.



For chemical equilibrium to be established the following needs to be present:

1. there must be a closed system
2. the concentrations of the reactants and products must remain constant
3. the rate of the forward reaction must be equal to the rate of the reverse reaction

Le Chatelier's Principle

When the equilibrium of a system is disturbed, the system tries to restore equilibrium by counteracting the factors causing the disturbance

There are three factors which can have an influence on equilibrium:

- Concentration
- Temperature
- Pressure

Effect of Concentration

Consider the reaction: $CoC\ell_4^{2-} + 6H_2O \rightleftharpoons Co(H_2O)_6^{2+} + 4C\ell^-$ $\Delta H < 0$

- Adding water:

  When more water is added to the solution, it makes more H_2O molecules available in the reaction. The equilibrium will shift to use up the extra H_2O, and we say that the forward reaction is favoured. This means the products concentration will increase and the solution will turn pink

- Adding hydrogen chloride ($HC\ell$)

  When hydrogen chloride is added, the $C\ell^-$ ion concentration will increase. This is also known as the **common ion effect**. The equilibrium will shift to use up the extra chloride ions. Thus the reverse reaction is favoured and the solution will turn blue

Effect of Temperature

Consider the reaction: $CoC\ell_4^{2-} + 6H_2O \rightleftharpoons Co(H_2O)_6^{2+} + 4C\ell^-$ $\Delta H < 0$

- Increase temperature:

  When the temperature of the system is increased, the system must try to decrease the temperature. Due to the reverse reaction being endothermic, the reverse reaction will be favoured and the reaction will turn blue.

- Decrease temperature

  When the temperature of the system is decreased, the system must try to increase the temperature. Due to the forward reaction being exothermic, the forward reaction will be favoured and the reaction will turn pink.

Effect of Pressure

Consider the reaction: $N_2O_4(g) \rightleftharpoons 2NO_2(g)$

- Increase pressure:

  When the pressure of the system is increased, the equilibrium will shift in order to decrease the pressure. The pressure will be decreased by favouring the reverse reaction due to there being less gas molecules.

- Decrease pressure

  When the pressure of the system is decreased, the equilibrium will shift in order to increase the pressure. The pressure will be increased by favouring the forward reaction due to there being more gas molecules.

Effect of a Catalyst

A catalyst does not affect the equilibrium, but it will increase the rate at which equilibrium is reached.

Demonstration

By changing factors like pressure, temperature and concentration the final concentrations of the reactants and products can be changed.

Aim

To investigation the effect of concentration and temperature on chemical equilibrium

Method

Prepare a solution of $CoC\ell_4^{2-}$ and place in several test tubes.

To one test tube add more water, and observe the colour change

To one test tube add hydrochloric acid and observe the colour change

Place the last test tube into an ice bath and then a beaker with warm water. Note the colour change.

The chemical reaction can be represented as follows:

$$CoC\ell_4^{2-} + 6H_2O \rightleftharpoons Co(H_2O)_6^{2+} + 4C\ell^- \qquad \Delta H < 0$$

The colour change in each experiment is due to the change in concentration of the products and reactions. When the reaction is blue in colour it is due to a high concentration of reactants and when the reaction is pink it is due to a high concentration of products.

Questions

Question 1

(Adapted from 2001 National Higher Grade Paper 2)

Copper is placed into nitric acid and a brown gas is collected in a gas syringe. The syringe is now sealed. The following equilibrium now exists in the syringe.

$$\underset{\text{brown}}{2NO_2(g)} \quad \rightleftharpoons \quad \underset{\text{colourless}}{N_2O_4(g)}$$

a.) When the syringe is placed in a beaker containing ice water, the colour of the gas in it becomes lighter. Is the forward reaction endothermic or exothermic?
b.) Explain your answer to question (a.) by making use of Le Chatelier's principle.
c.) How does the amount of NO_2 in the syringe change when the plunger is pushed in and equilibrium is re-established? (Write only increases, decreases or stays the same)

Question 2

(Adapted from 2001 IEB Higher Grade Paper 2)

Cobalt chloride crystals are dissolved in a beaker containing ethanol and then a few drops of water are added. After a period of time, the reaction reaches equilibrium as follows:

$$\underset{\text{blue}}{CoC\ell_4^{2-} + 6H_2O} \rightleftharpoons \underset{\text{pink}}{Co(H_2O)_6^{2+} + 4C\ell^-}$$

The solution, which is now just blue, is poured into three test tubes. State, in each case, what colour changes will be observed (if any) if the following are added in turn to each test tube:

a.) 1 cm^3 of distilled water
b.) A few crystals of sodium chloride
c.) The addition of hydrochloric acid to the third test causes the solution to turn pink. Explain why this occurs.

Question 3

(Adapted from 1999 Western Cape Higher Grade Paper 2)

0,1 mole of hydrogen gas (H_2) and an excess of sulphur powder are sealed in a 1,0 dm^3 container at a temperature of 90°C. The following equilibrium is established:

$$H_2(g) + S(s) \rightleftharpoons H_2S(g)$$

How will the concentration of the $H_2S(g)$ be affected if (Write only increases, decreases or stays the same)

a.) more sulphur is used?
b.) a smaller container is used?
c.) a suitable catalyst is added?
d.) more hydrogen is used?

Links

1. Demonstration of stimulated chemical equilibrium: http://www.youtube.com/watch?v=C5jDmG4nVV8
2. The concept of equilibrium: http://www.youtube.com/watch?v=LBjRDF4XV8Q
3. Chemical equilibrium and the equilibrium constant: http://www.youtube.com/watch?v=CGJjjm7-DJ4